

Chapter-11

Lesson-1

I – Kinetic Molecular Theory

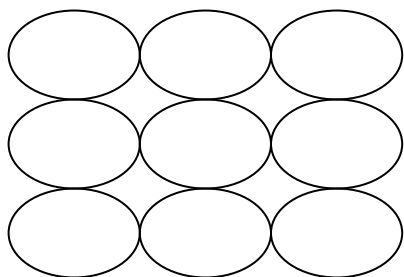
- The kinetic molecular theory was originally developed in the 19th century to explain the behavior of gases.
- However, the model also helps us to understand the behavior of liquids and solids as well.
- The table below summarizes the characteristics of gases, liquids, and solids and both the macroscopic and the microscopic levels.

	Gases	Liquids	Solids
Macroscale Level	<ul style="list-style-type: none">• Gases have an indefinite shape & an indefinite volume• Gases are very compressible.• Gases flow very readily.	<ul style="list-style-type: none">• Liquids have an indefinite shape & a definite volume• Liquids are incompressible.• Liquids flow readily.	<ul style="list-style-type: none">• Solids have a definite shape & a definite volume• Solids are incompressible.• Solids do not flow.
Mircoscale (molecular) Level	<ul style="list-style-type: none">• Particles are far apart.• Particles have complete freedom of motion.• Particles move about randomly.• Particles have negligible attractions for each other.	<ul style="list-style-type: none">• Particles are close together.• Particles can move about with restricted motion.• Particles move about randomly.• Particles have significant attractions for each other.	<ul style="list-style-type: none">• Particles are very closely packed in an ordered arrangement.• Particles can't move about.• Particles can vibrate in place.• Particles have very strong attractions for each other.

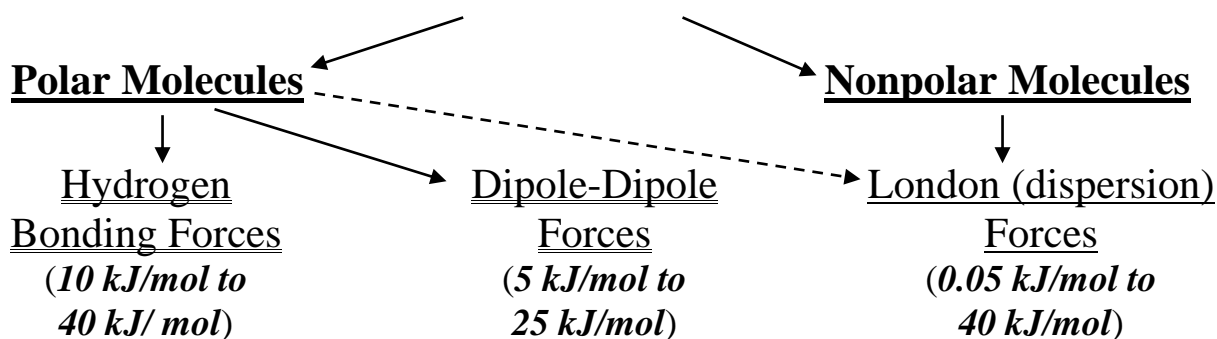
- In **true crystalline solids**, the unit particles are found in a *regular repeating internal arrangement* throughout the entire sample.
- The total three-dimensional structural pattern created by the unit particles in a true solid is called the **crystal lattice**. *True solids melt at a single temperature.*
- In **amorphous solids**, the unit particles lack a regular, long range pattern. *Amorphous solids melt over a range of temperatures.*

II - Intermolecular (van der Waal's) Forces (IMF's)

- All molecules experience forces of attraction for each other in all phases.
- These intermolecular forces of attraction (IMF's) are collectively known as **van der Waal's forces**.



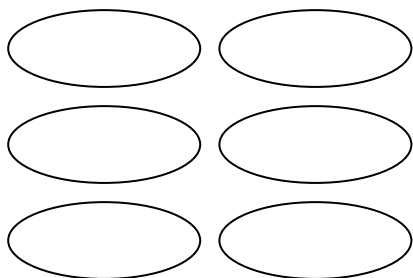
- The basic particles are individual molecules.
- Van der Waals forces are weak forces (*only 0.05 kJ/mol to 40 kJ/mol of energy is needed to break the van der Waals forces*)



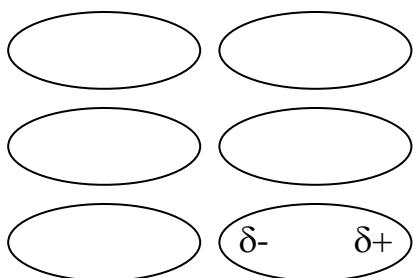
- The larger the mass of the molecules, the stronger the intermolecular forces of attraction (IMF's) generally are.
- Gram-for-gram, polar molecules have stronger IMF's than nonpolar molecules.
- Gram-for-gram, less compact molecules have stronger IMF's than more compact molecules. [**POLARIZABILITY**]
- Molecular substances with stronger IMF's have high melting and boiling points than molecular substances with weaker IMF's.

A) London (Dispersion) Forces of Attraction

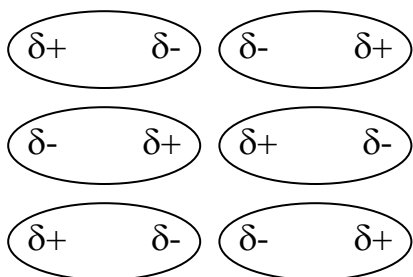
- All molecules experience London dispersion forces but their effect is only significant with nonpolar molecules.
- **London dispersion forces** can be weak (0.05 kJ/mol) to relatively strong (40 kJ/mol) depending on the size and polarizability of the attracting molecules.



Electrons are usually spread out evenly around a nonpolar molecule (that's why it's nonpolar).

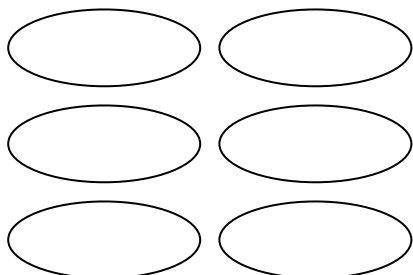


However, due to their random movement, electrons will, once in a while, be unevenly distributed around the molecule (OOPS)! This temporary uneven arrangement creates a ***temporary dipole***.



The temporary dipole will induce temporary polarization in all the neighboring molecules.

London dispersion Forces!

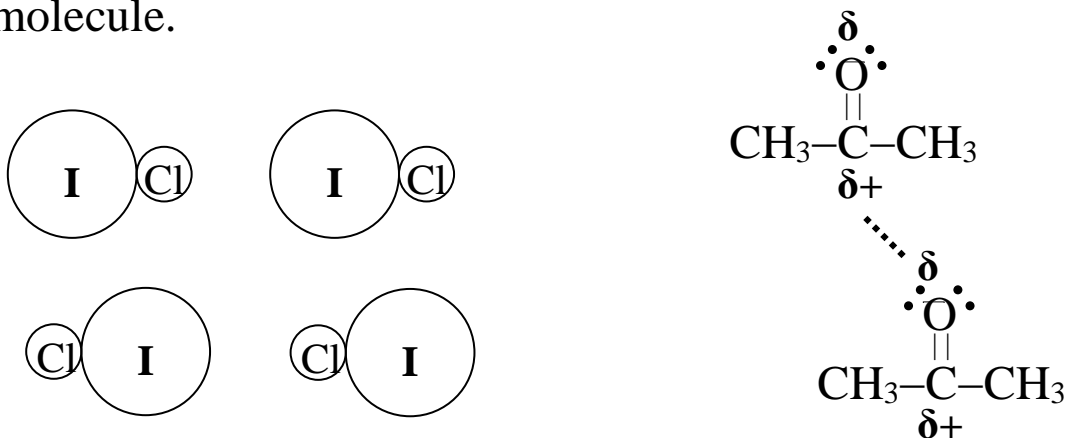


Of course, this arrangement only lasts for a fraction of a second because the electrons will be randomly redistributed over the molecules.

This process continually reoccurs in nonpolar molecular solids.

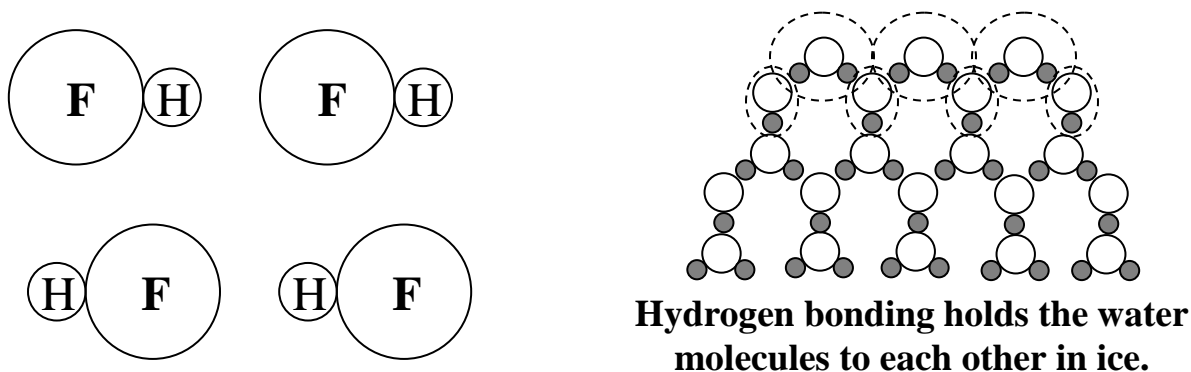
B) Dipole-Dipole Forces of Attraction

Polar molecules will line up so that the negative end of one polar molecule is adjacent to the positive end of a neighboring molecule.

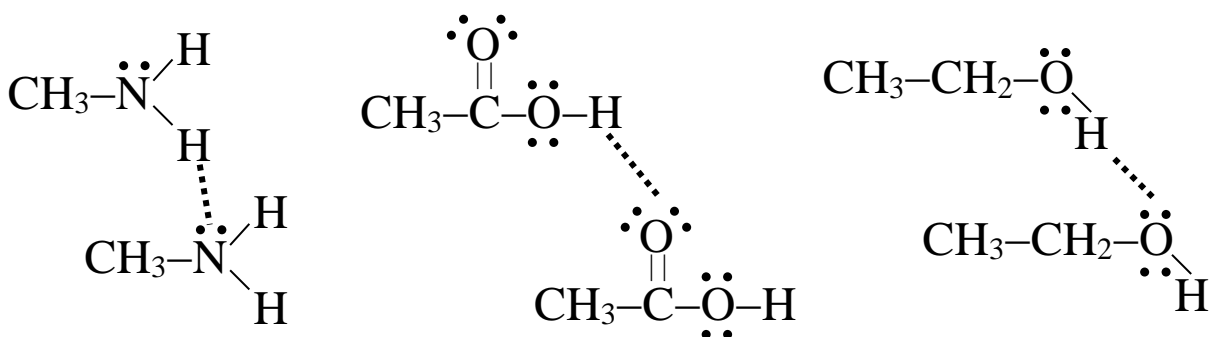


The attractive force between the oppositely charged ends of polar molecules is called a **dipole-dipole force**.

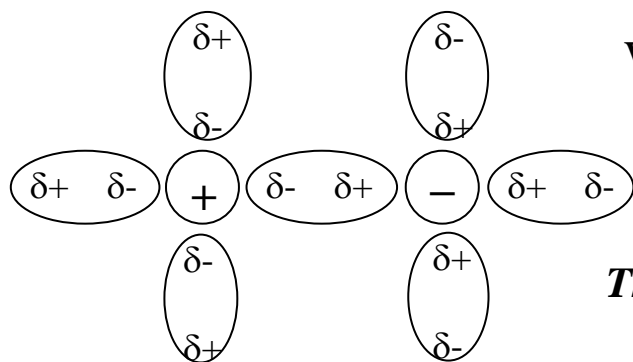
C) Hydrogen Bonding Forces of Attraction



Hydrogen bonding occurs when the nitrogen, oxygen, or fluorine atom from one molecule attracts a hydrogen atom from a neighboring molecule that is in turn bonded to a nitrogen, oxygen, or fluorine atom.



D) Ion-Dipole Forces of Attraction



When an ion and a nearby polar molecule attract each other an **ion-dipole force** is formed.

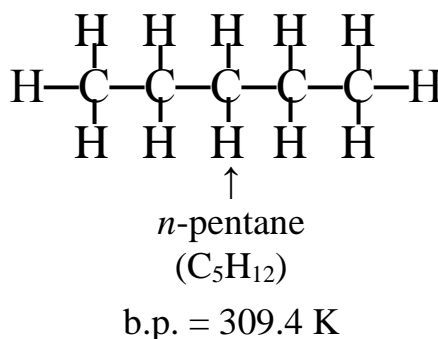
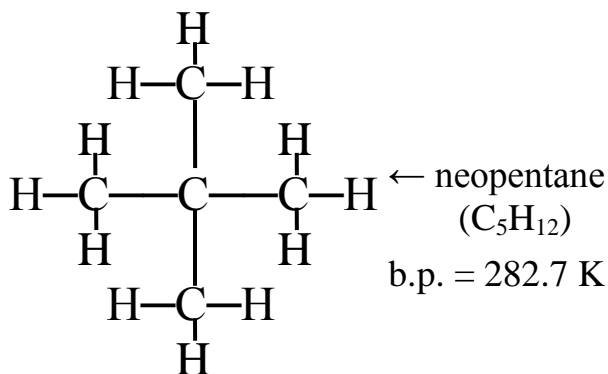
These are the forces that result when salts dissolve in water.

Q1: Explain why ammonia, water, and hydrogen fluoride have the highest boiling point in their respective group even though they have the lowest molar mass in their respective group?

Group 15 Hydrides		Group 16 Hydrides		Group 17 Hydrides	
Substance	b.p.	Substance	b.p.	Substance	b.p.
NH ₃	-33°C	H ₂ O	100°C	HF	20°C
PH ₃	-85°C	H ₂ S	-60°C	HCl	-85°C
AsH ₃	-55°C	H ₂ Se	-45°C	HBr	-67°C
SbH ₃	-20°C	H ₂ Te	-5°C	HI	-35°C

A1: Samples of ammonia, water, and hydrogen fluoride are primarily held in the liquid phase by hydrogen bonding whereas the other compounds in the table are primarily held in the liquid phase by weaker dipole-dipole forces.

Q2: Explain why neopentane has a lower boiling point than the pentane even though both have the same molecular mass?



A2: Neopentane is composed of more compact molecules than pentane which makes neopentane molecules less polarizable.

III Viscosity and Surface Tension

A) Viscosity:

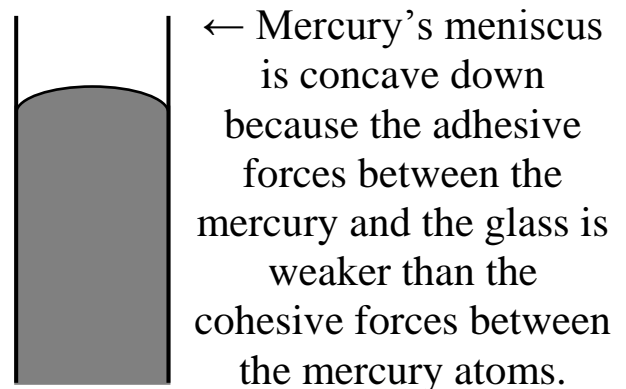
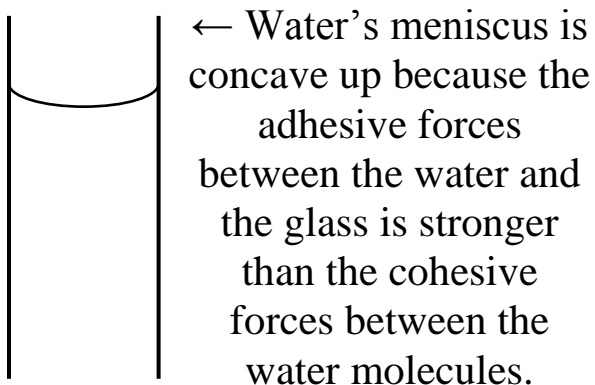
- Liquids flow readily. However, some liquids flow more readily than others. The resistance of a liquid to flow is called **viscosity**.
- A liquid's viscosity depends on the attractive forces between the molecules.
- A liquid's viscosity also depends on the ability of the molecules to become entangled with each other.
- The higher a liquid's viscosity value, the more slowly it will flow.

A motor oil with a 40 viscosity rating is “thicker” than a motor oil with a 30 viscosity rating.

- A liquid's viscosity increases with a decrease in temperature.

B) Surface Tension:

- All liquids have **surface tension** which is caused by the cohesive attractive forces between the unit particles at the surface of the liquid and the rest of the unit particles in the sample.
- Unit particles at the surface of a liquid experience a net attractive force toward the center of the sample. Surface tension is a measure of the energy to disrupt this attractive force.
- **Cohesion** is the attractive force between a substance's unit particles. **Adhesion** is the attractive force between a substance's unit particles and another surface.

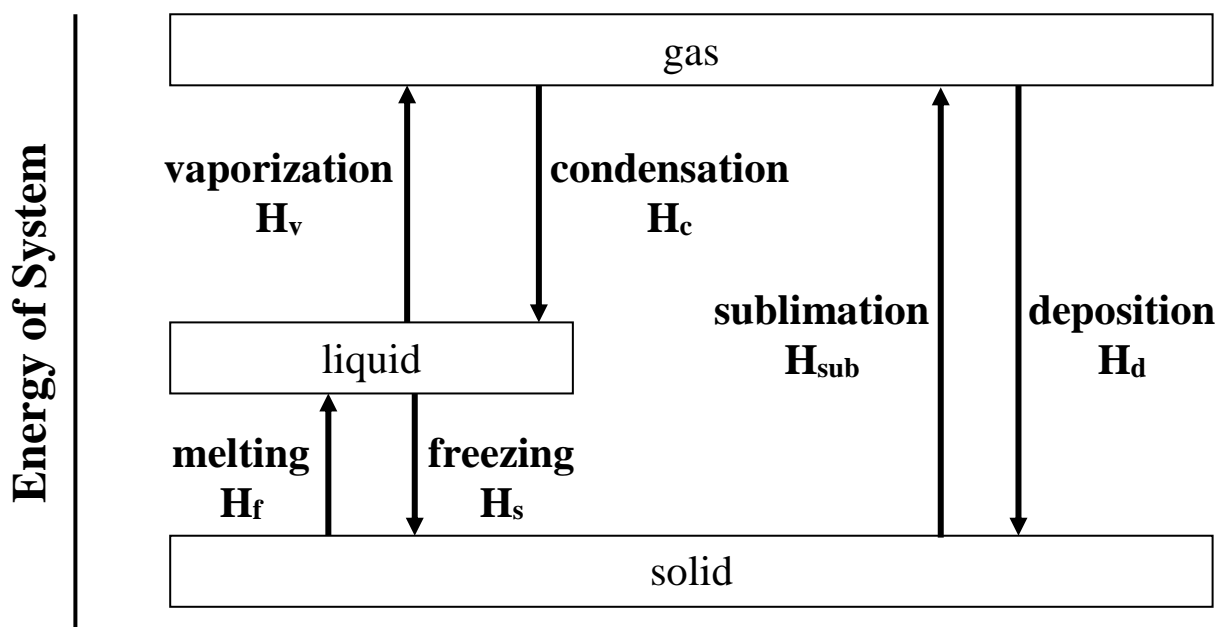


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Lesson-2

IV – Changes of State

A) Energy Changes Accompanying Phase Changes:



- For a pure substance ...

$$H_f = -H_s \quad \& \quad H_v = -H_c \quad \& \quad H_{sub} = -H_d$$

- For water ...

$$H_f = 334 \text{ J/g} \quad \text{or} \quad 6.01 \text{ kJ/mol}$$

$$H_v = 2260 \text{ J/g} \quad \text{or} \quad 40.7 \text{ kJ/mol}$$

$$H_{sub} \approx 2590 \text{ J/g} \quad \text{or} \quad 46.7 \text{ kJ/mol}$$

$\frac{\text{J}}{\text{g}}$
↓

$\frac{\text{kJ}}{\text{mol}}$
↓

- The change in heat for a phase change is $q = m\Delta H$ or $q = n\Delta H$

Q1: How much heat would be released if a 55.0 gallon drum filled with water (208197g) at 0.0°C were to completely freeze?

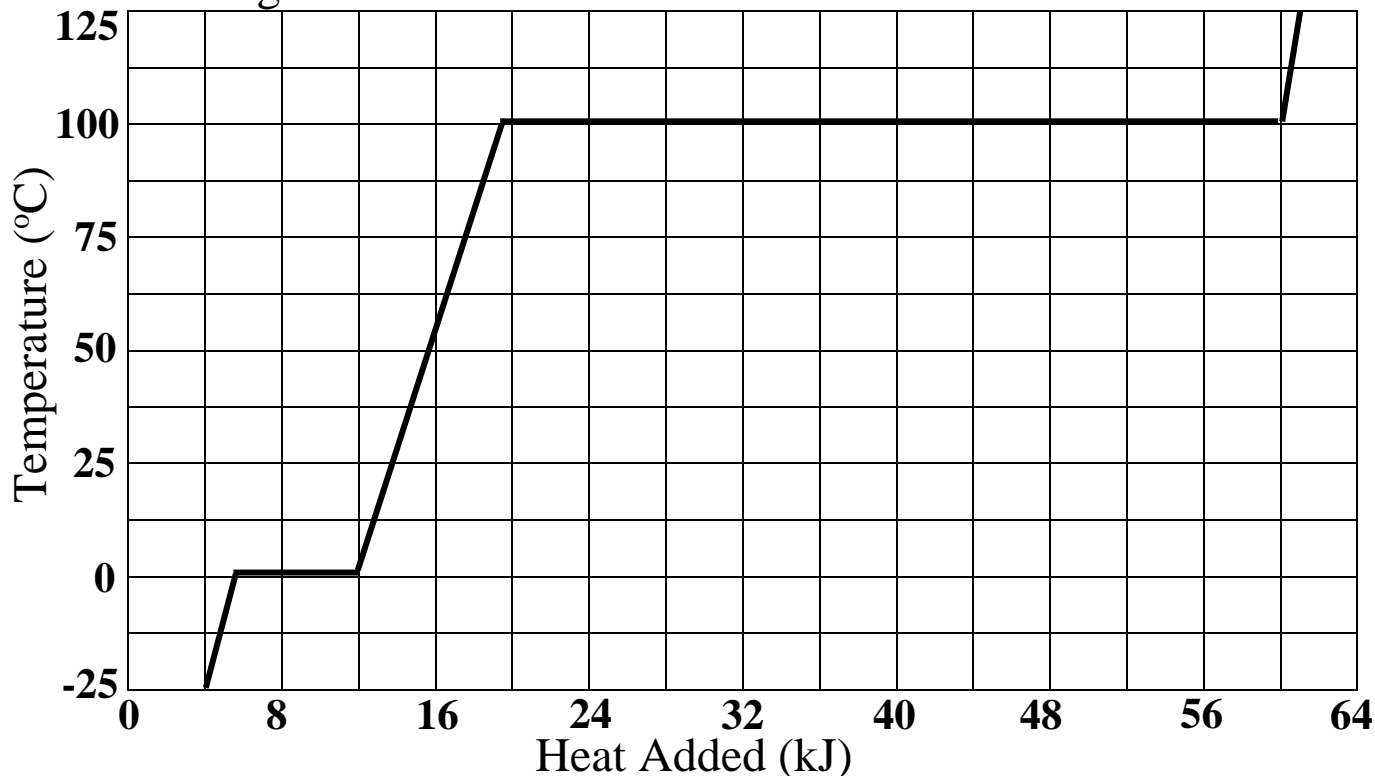
A1: $q = mH_s$

$$q = (208197 \text{ g})(-334 \text{ J/g})$$

$$q = -6.95 \times 10^7 \text{ J of heat} \leftarrow \text{an exothermic change!}$$

B) Heating Curves:

Heating Curve for 1.00 Mole of Water



- The change in heat for a temperature change is ...

$$q = mc\Delta T \quad \text{or} \quad q = nC\Delta T$$

- The specific heat for ice is ...

$$c = 2.09 \text{ J/g}^\circ\text{C} \quad \text{or} \quad C = 37.7 \text{ J/mol}^\circ\text{C}$$

- The specific heat for water is ...

$$c = 4.18 \text{ J/g}^\circ\text{C} \quad \text{or} \quad C = 75.3 \text{ J/mol}^\circ\text{C}$$

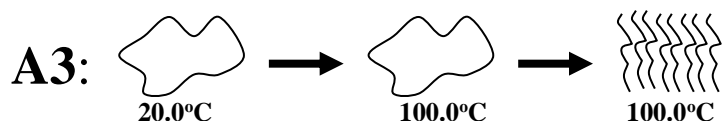
- The specific heat for steam is ...

$$c = 1.84 \text{ J/g}^\circ\text{C} \quad \text{or} \quad C = 33.2 \text{ J/mol}^\circ\text{C}$$

Q2: Explain why the heat of fusion of any substance is always lower than its heat of vaporization?

A2: Melting requires that the molecules do not separate from each other whereas vaporization requires that the molecules widely separate from each other.

Q3: How much heat would be needed to raise the temperature of 2000.0 g of water at 20.0°C to steam at 100.0°C?



$$q = mc_{\ell}\Delta T + mH_v$$

$$q = (2000.0 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(80.0^\circ\text{C}) + (2000.0 \text{ g})(2260 \text{ J/g})$$

$$q = 5190000 \text{ J of heat added}$$

Q4: For many centuries in hot arid climates, water has been cooled inside porous clay evaporation jars. How many grams of water can be cooled from 35°C to 22°C by the evaporation of only 50.0 g of water? [at 35°C, the $H_{\text{evap}} = 2400 \text{ J/g}$]

A4: $q_{\text{evap}} = mH_{\text{evap}}$

$$q_{\text{evap}} = (50.0 \text{ g})(2400 \text{ J/g})$$

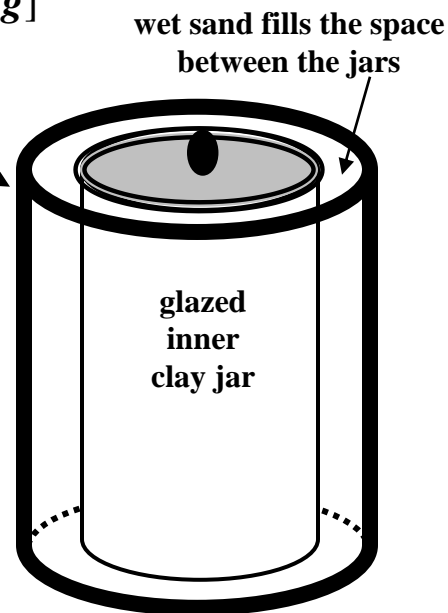
$$q_{\text{evap}} = 120000 \text{ J of heat removed by evaporation}$$

$$q_{\text{lost by H}_2\text{O}} = m_{\text{of H}_2\text{O}} \times c_{\text{of H}_2\text{O}} \times \Delta T_{\text{of H}_2\text{O}}$$

$$120000 \text{ J} = (X)(4.18 \text{ J/g}^\circ\text{C})(13^\circ\text{C})$$

$$X = 2200 \text{ g} \rightarrow 2.2 \text{ L of cooled water}$$

unglazed
porous outer
clay jar



Q5: Name the phase transition in each of the following situations and indicate whether it is exothermic or endothermic:

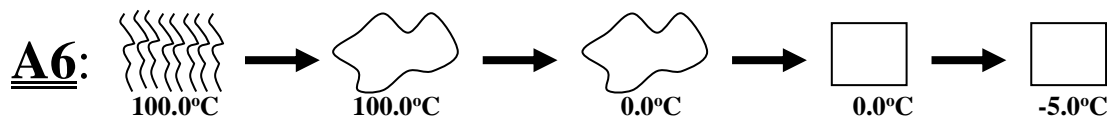
(i) Wet clothes dry on a hot day. (ii) Frost appears on a window on a cold day. (iii) An iced tea glass “sweats” on a hot day.

A5: (i) evaporation; endothermic

(ii) deposition; exothermic

(iii) condensation; exothermic

Q6: When frost threatens the orange groves in Florida, the farmers set out steam pots under the trees. How much heat can be released if 2.00×10^6 g of steam at 100.0°C that is released in an orange grove whose temperature is -5.0°C ?

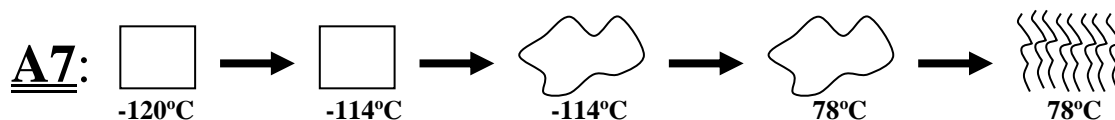


$$q = mH_c + mc_\ell\Delta T + mH_s + mc_s\Delta T$$

$$q = (2.00 \times 10^6 \text{ g})(2260 \text{ J/g}) + (2.00 \times 10^6)(4.18 \text{ J/g}^\circ\text{C})(100.0^\circ\text{C}) + \\ (2.00 \times 10^6 \text{ g})(334 \text{ J/g}) + (2.00 \times 10^6 \text{ g})(2.09 \text{ J/g}^\circ\text{C})(5.0^\circ\text{C})$$

$$q = 6.04 \times 10^9 \text{ J of heat released}$$

Q7: Ethanol melts at -114°C and boils at 78°C . Its H_f is 109 J/g , and its H_v is 838 J/g . The specific heat of solid ethanol is $0.97 \text{ J/g}^\circ\text{C}$ and the specific heat of liquid ethanol is $2.30 \text{ J/g}^\circ\text{C}$. How much heat is required to convert 75.0 g of solid ethanol at -120°C to gaseous ethanol at 78°C ?



$$q = mc_s\Delta T + mH_f + mc_\ell\Delta T + mH_v$$

$$q = (75.0\text{g})(0.97 \text{ J/g}^\circ\text{C})(6.0^\circ\text{C}) + (75.0\text{g})(109 \text{ J/g}) + \\ (75.0\text{g})(2.30 \text{ J/g}^\circ\text{C})(192.0^\circ\text{C}) + (75.0\text{g})(838 \text{ J/g}^\circ\text{C})$$

$$q = 105000 \text{ J of heat added}$$

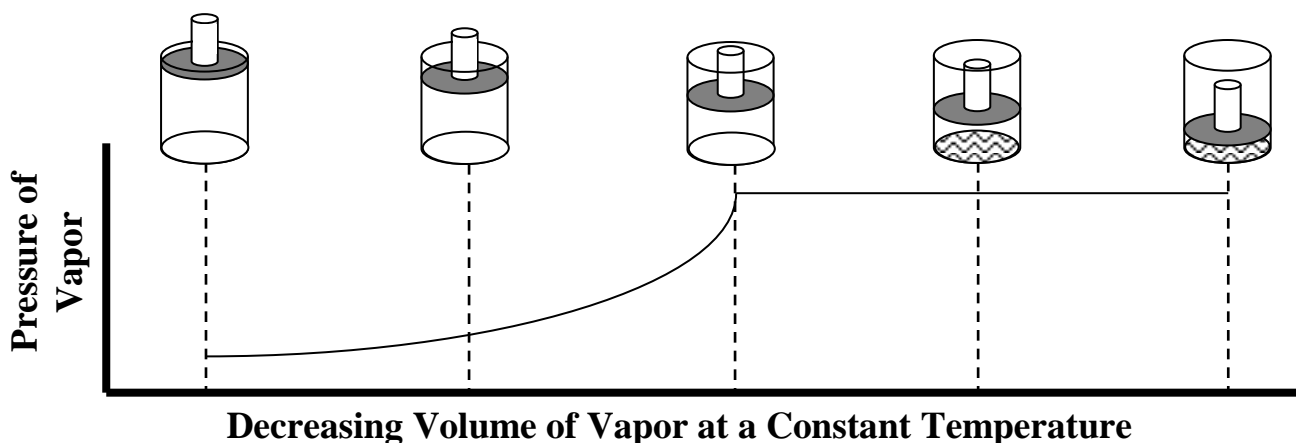
Chapter-11

Lesson-3

V – Vapor Pressure and Boiling Point

A) Vapor Pressure:

- All liquids continually evaporate at their upper surface.
- In a closed container, a dynamic equilibrium is eventually established between the opposing processes of evaporation and condensation. *liquid \rightleftharpoons vapor*
- The pressure of vapor in equilibrium with a liquid is called the **vapor pressure** and is the maximum allowed pressure.
- As long as both liquid and vapor are present, the pressure exerted by the vapor (*the vapor pressure*) is independent of the volume of the container but it is a function of the liquid's temperature.



Q1: If 4.00 L of diethyl ether vapor at 700. mmHg and 40.0°C is compressed to 2.50 L at 40.0°C, then will condensation occur?
The vapor pressure for diethyl ether is 1000. mmHg at 40.0°C

A1: $(P_1)(V_1) = (P_2)(V_2)$

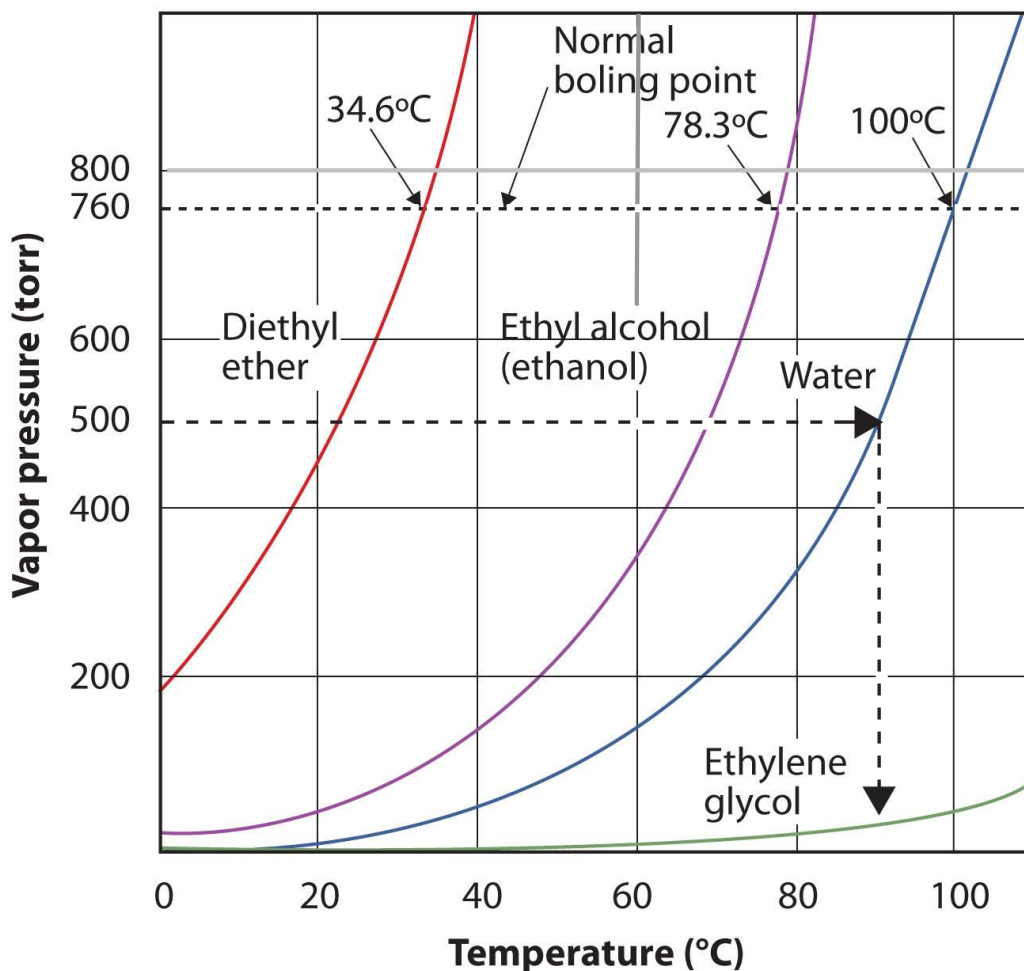
$$(700.)(4.00) = (x)(2.50)$$

$x = 1120 \text{ mmHg}$ \leftarrow *the pressure if diethyl ether were an ideal gas*

YES *some condensation will occur.*

B) Vapor Pressure versus Temperature:

- A liquid's vapor pressure increases as its temperature increases and decreases as its temperature decreases.



Q2: If the pressure of acetone vapor in a rigid container at 40.0°C is 250 Torr, then using ideal gas law what would be the pressure at 30.0°C (*VP = 270 Torr*)? And at 20.0°C (*VP = 205 Torr*)?

A2:

At 30.0°C

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{250}{313} = \frac{x}{303}$$

$$x = 242 \text{ Torr (ideal)}$$

$$P = 242 \text{ Torr}$$

At 20.0°C

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\frac{250}{313} = \frac{x}{293}$$

$$x = 234 \text{ Torr (ideal)}$$

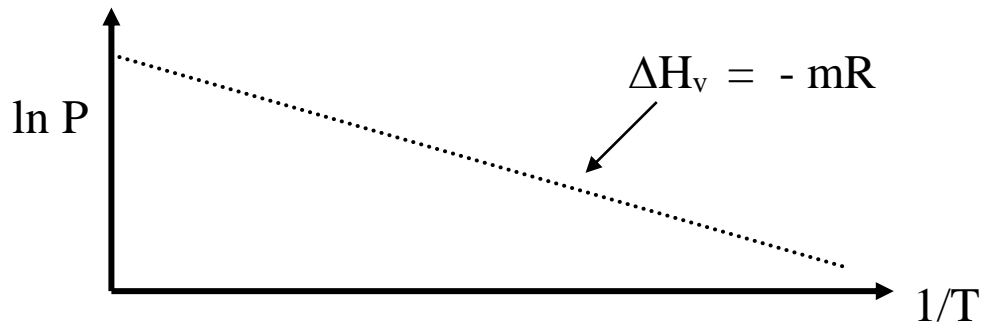
$$P = 205 \text{ Torr (condensation)}$$

- The nonlinear relationship shown in the **Vapor Pressure vs Temperature** graph can be expressed as a linear relationship between "ln P" and "1/T":

$$\ln P = \frac{-\Delta H_v}{R} \left(\frac{1}{T} \right) + C$$

$$y = m \cdot x + b$$

- This is the **Clausius-Clapeyron equation**.



- A two-point version of the equation can be used to determine a liquid's heat of vaporization (ΔH_v).

$$\ln \frac{P_1}{P_2} = \frac{-\Delta H_v}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

Q3: If the vapor pressure for water is 760.0 mmHg at 100.0°C, then calculate the vapor pressure at 37.0°C. ($H_v = 40.7$ kJ/mol)

A3: $\ln \frac{P_1}{760.0} = - \frac{40700}{8.314} \left[\frac{1}{310.0} - \frac{1}{373.0} \right]$

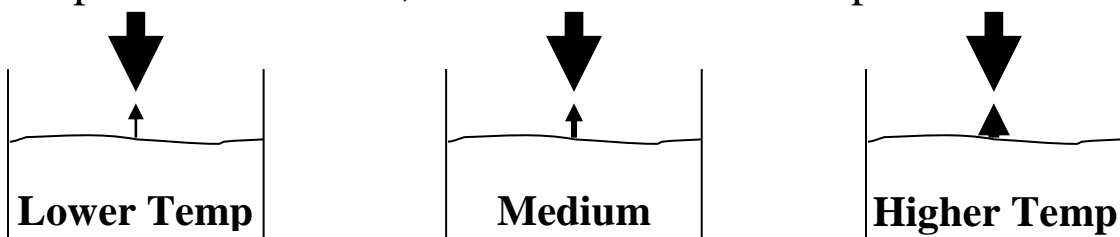
$$\ln P_1 - 6.63 = - 2.67$$

$$\ln P_1 = 3.96$$

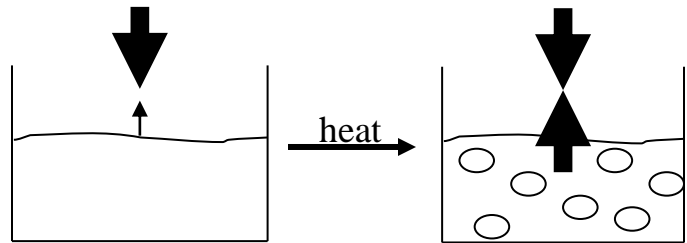
$$P_1 = 52.5 \text{ mmHg}$$

C) Vapor Pressure and Boiling Point Temperature:

- A liquid evaporates at any temperature. As the liquid's temperature increase, so does the rate of evaporation.



- A liquid can only boil when its vapor pressure is equal to the prevailing air pressure.



- If the air pressure is higher than normal then the liquid will boil at a temperature above its normal boiling point and vice versa.

D) Critical Temperature and Pressure

- There is a temperature, called the critical temperature, above which the liquid phase of a pure substance cannot exist because the gas cannot be condensed regardless of the pressure applied.
- The pressure required to condense a gas at its critical temperature is called the critical pressure.
- The critical pressure is the maximum possible vapor pressure for a liquid.

"Permanent Gases"	Critical Temp (°C)	"Condensable Gases"	Critical Temp (°C)
Helium	- 268	Carbon dioxide	31
Hydrogen	- 240	Ethane	32
Nitrogen	- 147	Propane	97
Argon	- 122	Ammonia	132
Oxygen	- 119	Chlorine	144
Methane	- 82	Sulfur dioxide	158

Q4: Calculate the boiling point of water at the summit of Mount Whitney, CA if the daily air pressure is only 500.0 torr.

A4: $\ln \frac{500.0}{760.0} = -\frac{40700}{8.314} \left[\frac{1}{T_1} - \frac{1}{373.0} \right]$

$$-0.4187 = -\frac{4900}{T_1} + 13.1$$

$$\frac{4900}{T_1} = \frac{13.5}{1}$$

$T_1 = 363 \text{ K} \rightarrow 90.^{\circ}\text{C}$

consult the vapor pressure curve two pages earlier!

Q5: If pure liquid nitric acid has a vapor of 47.9 Torr at 20.0°C and 670. Torr at 80.0°C, then (i) calculate its heat of vaporization in **kJ/mol** and (ii) its normal boiling point in °C using the vapor pressure of 208 Torr at 50.0°C as one of the reference points.

$$\underline{\underline{\text{A5}}}: \text{(i)} \quad \ln \frac{47.9}{670.} = - \frac{X}{8.314} \left[\frac{1}{293.0} - \frac{1}{353.0} \right]$$

$$- 2.63 = - \frac{X}{8.31} \left[0.000580 \right]$$

$$X = 37700 \text{ J/mol} \rightarrow \boxed{37.7 \text{ kJ/mol}}$$

$$\text{(ii)} \quad \ln \frac{760.}{208} = - \frac{37700}{8.314} \left[\frac{1}{T_1} - \frac{1}{323.0} \right]$$

$$1.30 = - \frac{4540}{T_1} + 14.0$$

$$\frac{4540}{T_1} = \frac{12.7}{1}$$

$$\boxed{T_1 = 357 \text{ K} \rightarrow 84^\circ\text{C}}$$

Q6: Acetone's heat of vaporization is 32.0 kJ/mol and its normal boiling point is 56.5°C. (i) Calculate its vapor pressure in **Torr** at 0.0°C. (ii) Sketch acetone's vapor pressure curve on the graph on the next page (*See question-2 for another set of data points*).

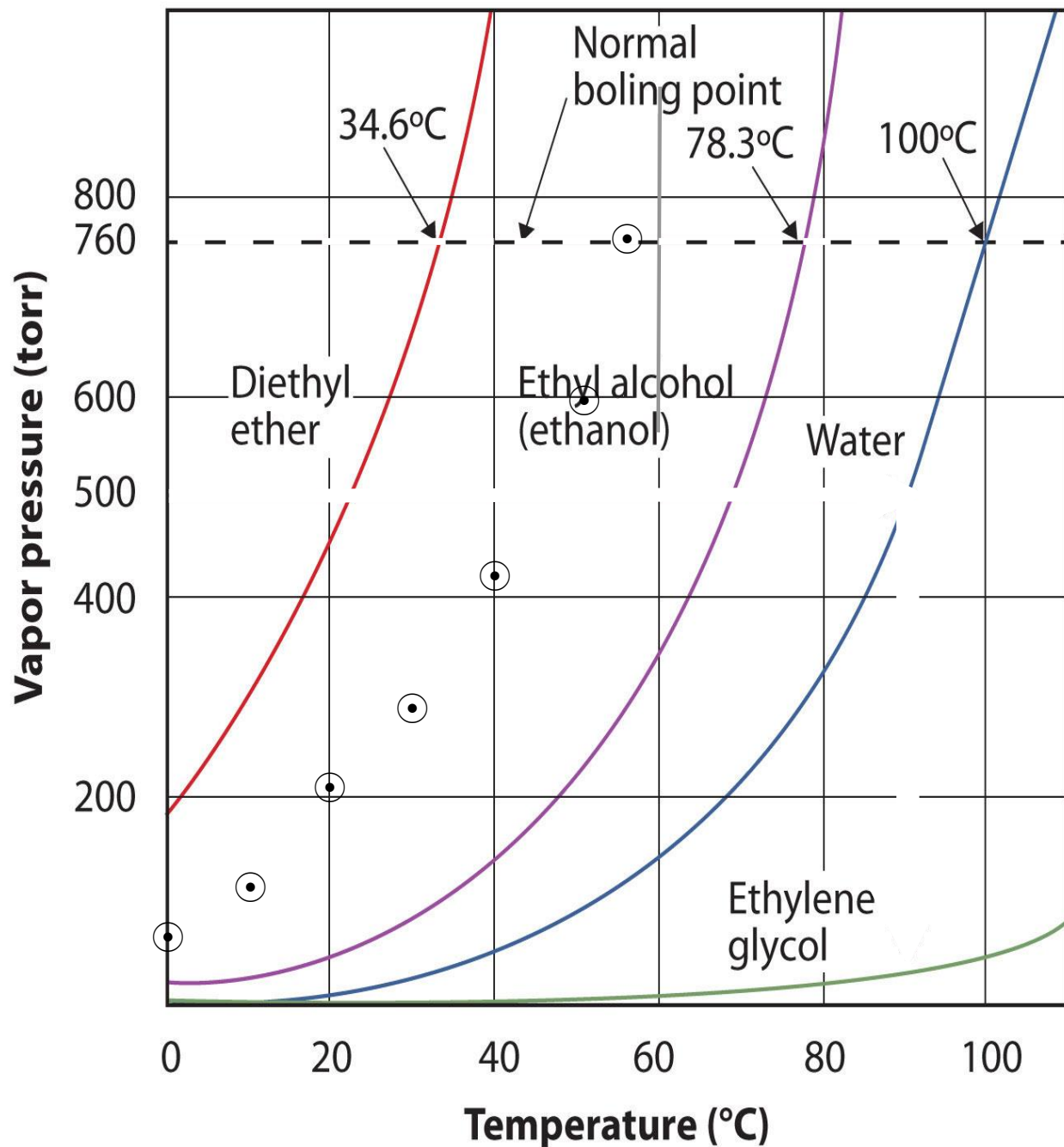
$$\underline{\underline{\text{A6}}}: \text{(i)} \quad \ln \frac{760.}{X} = - \frac{32000}{8.314} \left[\frac{1}{329.5} - \frac{1}{273.0} \right]$$

$$6.63 - \ln X = 2.42$$

$$\ln X = 4.21$$

$$X = 67.4 \text{ Torr}$$

(ii)



Q7: In terms of kinetic molecular theory, why is it impossible for a substance to exist as a liquid above its critical temperature?

A7: At the critical temperature, all molecules have kinetic energies greater than any intermolecular forces and so a liquid cannot form.

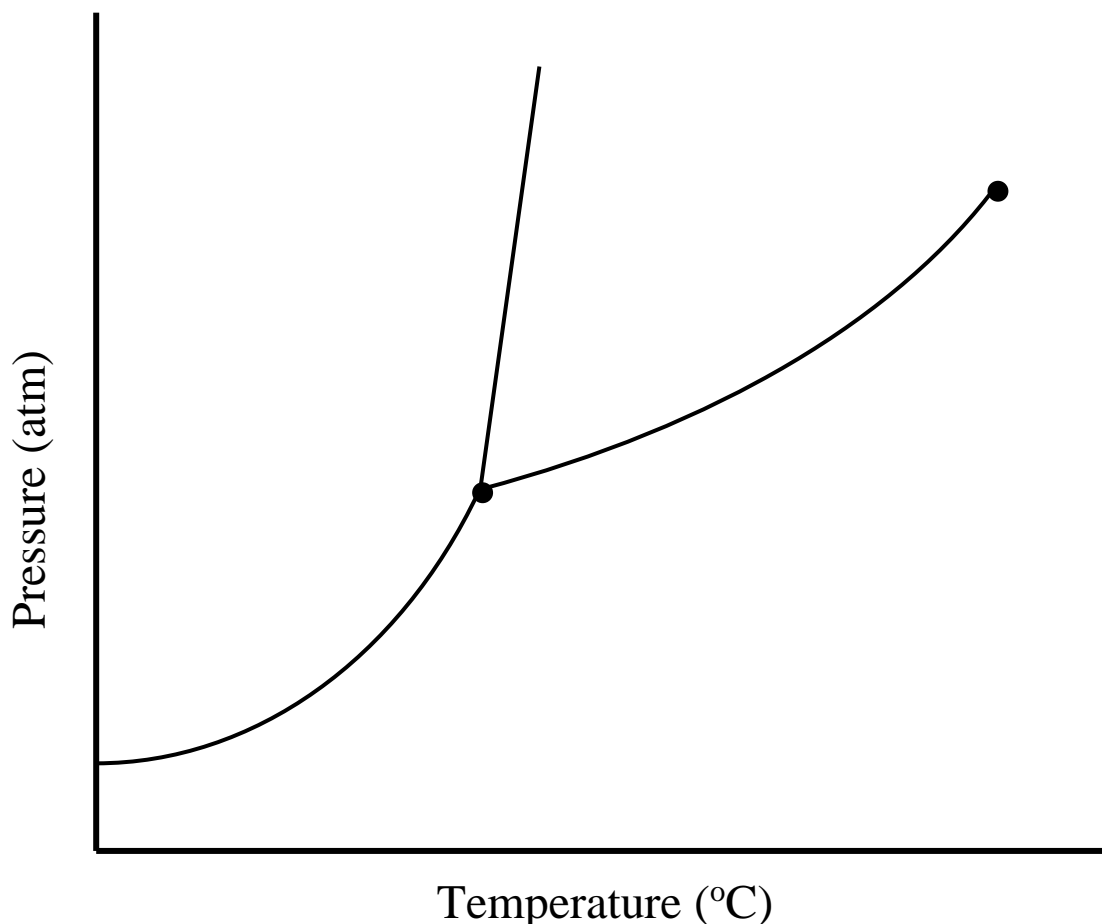
Chapter-11

Lesson-4

VI – Phase Diagrams

- On a phase diagram, the temperature and pressure point at which all three phases of a pure substance can simultaneously exist is called the **triple point**.
- On a phase diagram, the critical temperature and critical pressure point of a pure substance is called the **critical point**.

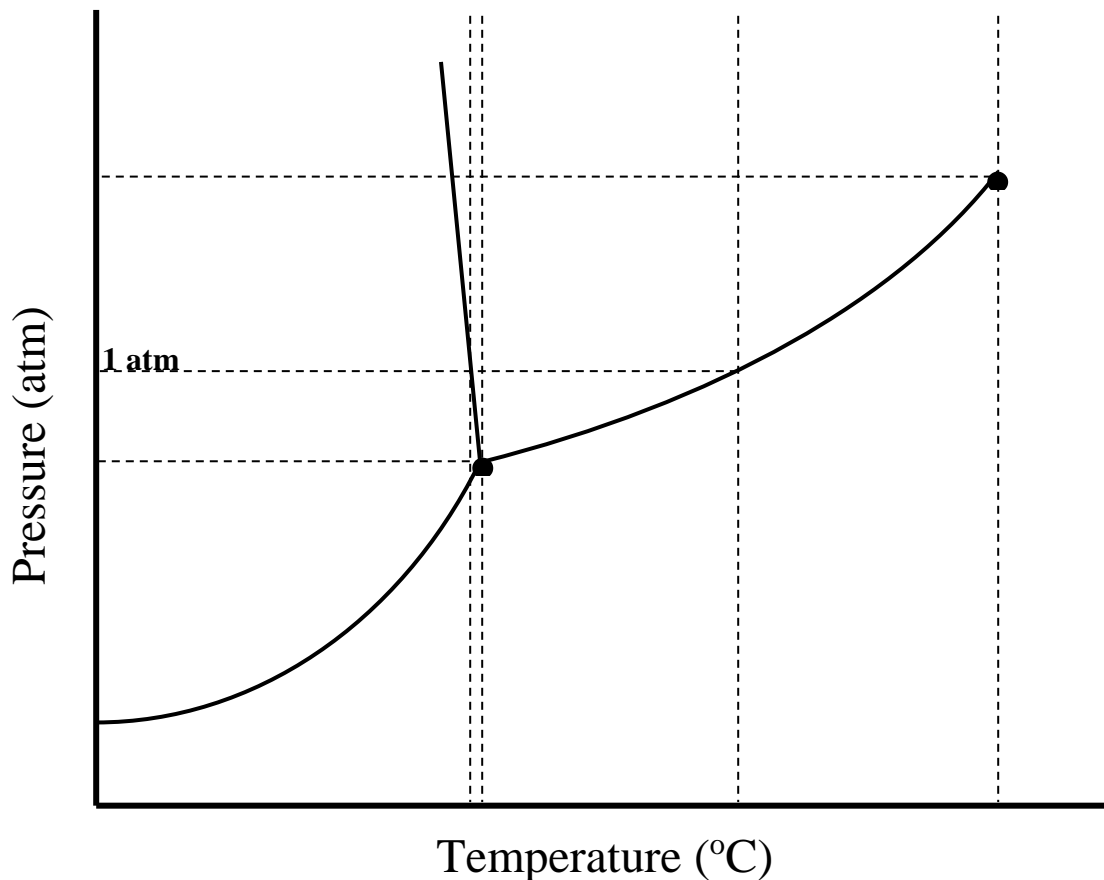
A) Solid-Liquid Line with Positive Slope:



Q1: A substance's triple point is at 60. mmHg and 85°C. In what phase is the substance at 30. mmHg and 105°C?

A1: Gas

B) Solid-Liquid Line with Negative Slope:



- The triple point for water occurs at 0.0098°C and 0.00603 atm.
- The critical point for water occurs at 374°C and 218 atm.
- The vapor pressures for ice and water at the triple point are equal (0.00603 atm).
- The curve from the triple point to the critical point is the vapor pressure curve!
- Thus, on a phase diagram the temperature point at which the pressure is 1.00 atm is the normal boiling point.

Q2: What is the slope of the “*solid ice – liquid water*” equilibrium line?

A2: $m = \frac{Y_2 - Y_1}{X_2 - X_1}$

$$m = \frac{1.00000 \text{ atm} - 0.00603 \text{ atm}}{0.0000^\circ\text{C} - 0.0098^\circ\text{C}}$$

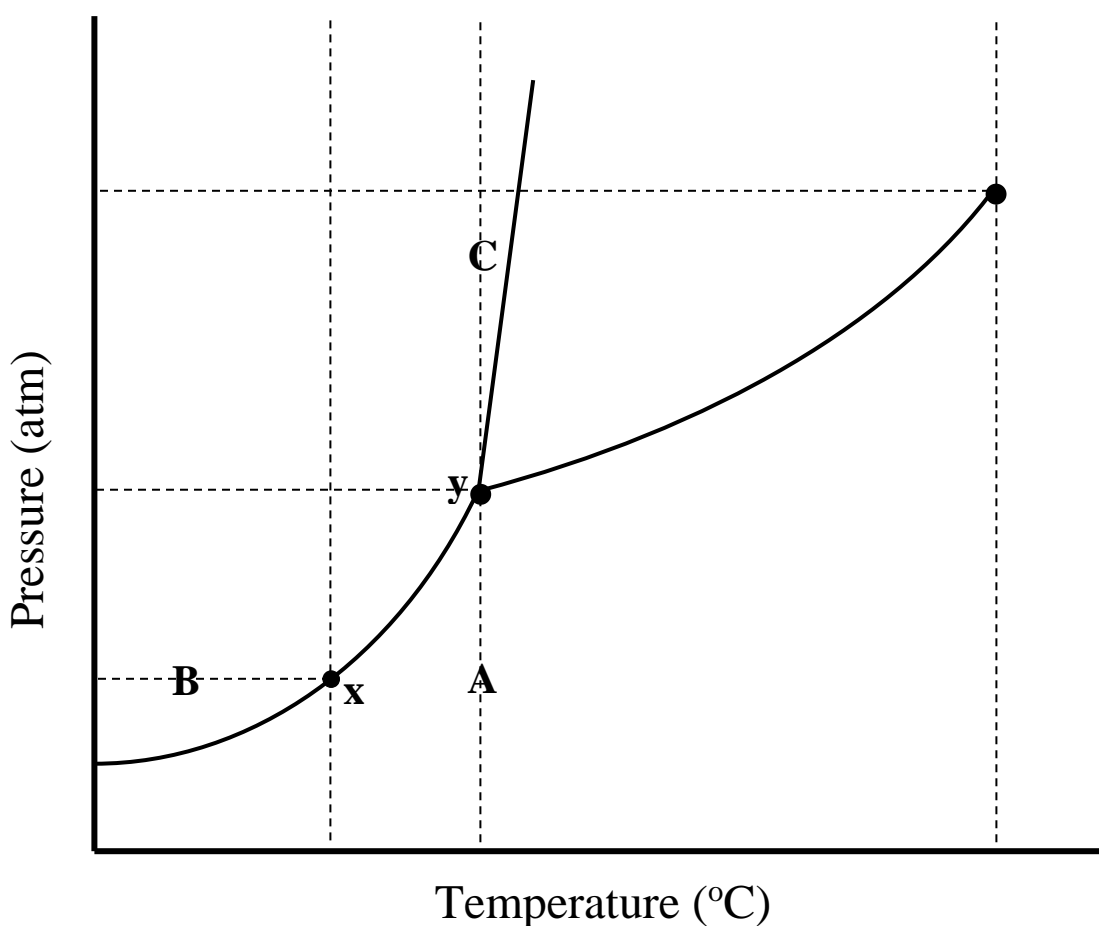
$$m = -100 \text{ atm}/^\circ\text{C} \quad \leftarrow \text{very steep slope indeed!}$$

Q3: What is the melting point of ice at 218 atm?

$$\text{A3: } \frac{-100 \text{ atm}}{1.00^\circ\text{C}} = \frac{218 - 1}{x - 0.00}$$

$$x = -2.17^\circ\text{C}$$

Q4: Describe the changes that CO_2 undergoes as the conditions change (a) from point A to point B at constant pressure and (b) from point A to point C at constant temperature.

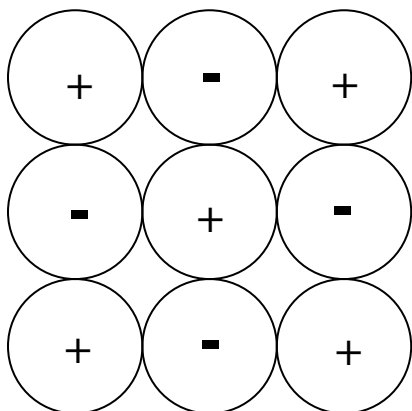


A4: (a) Gaseous CO_2 at point-A is cooled to point-x at which point it begins to solidify (by deposition) to dry ice. The dry ice is then further cooled to point-B.

(b) Gaseous CO_2 at point-A is pressurized to point-y at which point it begins to solidify (by deposition) to dry ice. The dry ice is then further pressurized to point-C.

VII - Classes of Solid

A) Ionic Solids:

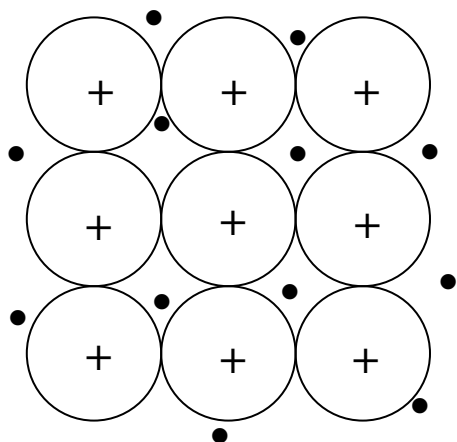


The basic particles in ionic solids are cations and anions.

The forces holding the ions in place are ionic bonds.

Ionic bonds are strong forces
(400 kJ/mol to 4000 kJ/mol)

B) Metallic Solids:

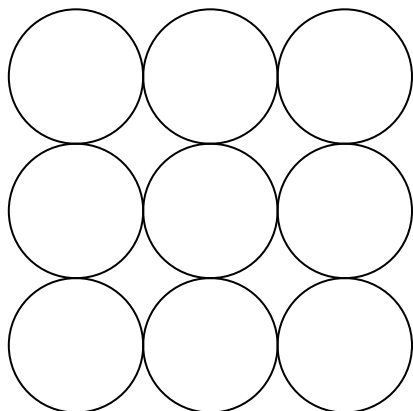


The basic particles in metallic solids are metallic cations.

The forces holding the cations in place are metallic bonds.

Metallic bonds are strong forces
(75 kJ/mol to 1000 kJ/mol)

C) Network Solids:

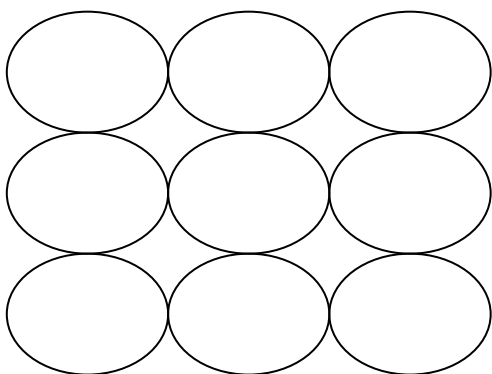


The basic particles in network solids are nonmetallic atoms.

The forces holding the atoms in place are covalent bonds.

Covalent bonds are strong forces
(150 kJ/mol to 1100 kJ/mol)

D) Molecular Solids:



The basic particles in molecular solids are individual molecules (*which may be very small or very large*).

The forces holding the molecules in place are van der Waal's forces.

Overall, van der Waal's forces are weak forces (**0.05 kJ/mol to 40 kJ/mol**)

Q5: Classify each of the following solids as ionic, metallic, network, or molecular: *Xe*, *SiO₂*, *Na*, *C₁₀H₈*, *C*, *NaCl*, *Al*, and *MgSO₄*

A5: Ionic: *NaCl*, *MgSO₄*

Metallic: *Na*, *Al*

Network: *SiO₂*, *C*

Molecular: *C₁₀H₈*, *Xe*

Q6: Describe the nature of the structural units in each of the following substances in the solid phase: *K₂O*, *O₂*, *K*, and *BN*

A6: *K₂O* is composed of K^{1+} ions and O^{2-} ions

O₂ is composed of *O₂* molecules

K is composed of K^{1+} ions surrounded by a mobile sea of e^{-} 's

BN is composed of B atoms and N atoms

Q7: Of the four general classes of solids, which ones

(a) are generally low boiling?

(b) are ductile and malleable?

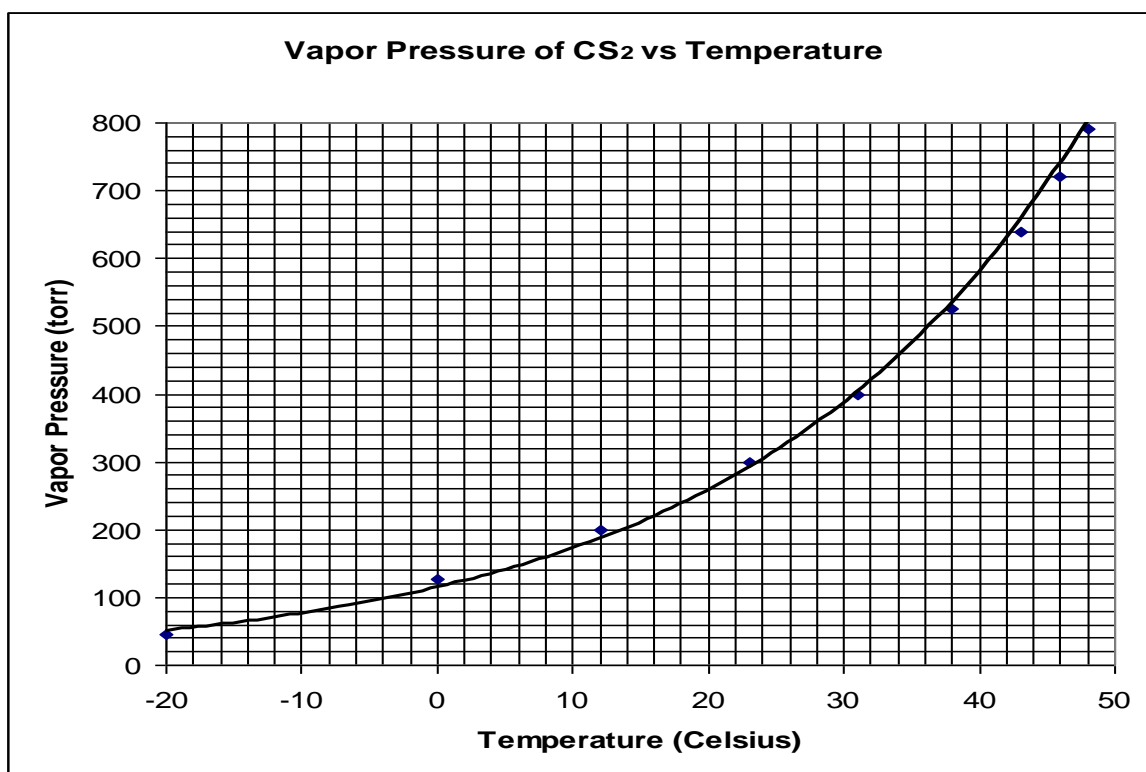
(c) are generally nonvolatile?

A7: (a) molecular solids

(b) metallic solids

(c) metallic solids, ionic solids, and network solids

Q8: Using the graph below, determine the normal boiling point of CS₂.



A8: $\approx 47^{\circ}\text{C}$

Q9: Covalent bonding occurs in both molecular and network solids. Why do these two kinds of solids differ so greatly in their hardness and boiling points?

A9: Weak intermolecular forces bind the molecules to each other in the lattice of molecular solid, so that relatively little energy is needed to disrupt these forces. Strong covalent bonds hold the atoms to each other in the lattice of network solids, so that a large amount of energy is needed to disrupt these forces.

Q10: For each of the following pairs of solid substances, predict which will have the higher melting point and indicate why?
(i) SiO_{2(s)} and CO_{2(s)} (ii) Ar_(s) and Xe_(s)

A10: (i) SiO_{2(s)} (*network solid*) > CO_{2(s)} (*molecular solid*)

(ii) Xe_(s) (*larger molecular solid*) > Ar_(s) (*smaller molecular solid*)
Larger molecules are more polarizable and have stronger IMF's.

Chapter-11

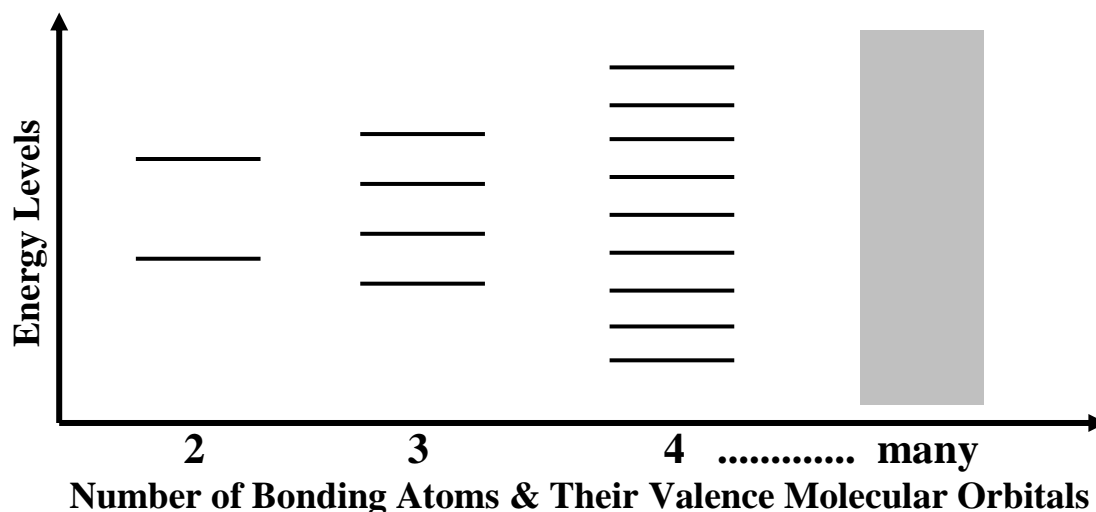
Lesson-5

VIII - Types of Solid Materials

A) Metals and Metallic Alloys:

1) Metals

- In the *electron-sea model* of metallic bonding, the metal is imagined as an array of cations held together by a mobile sea of electrons.
 - ☛ This model is adequately able to explain the conductivity as well as the malleability and ductility of metals.
 - ☛ This model is NOT able to adequately explain the trends in bonding strength, heat of fusion, hardness, melting point, or luster.
- The *molecular-orbital model* of bonding is able to explain all of these properties.
 - ☛ Atomic orbitals combine to form molecular orbitals as atoms bond – even when they bond metallicity.



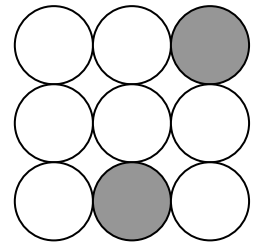
- ☛ In metals, the valence molecular orbitals form a continuous band of allowed energy levels called the *energy band*.
- ☛ Valence electrons in filled MOs of lower energy are easily able to move to empty MOs of higher energy.

- The ability of electrons to absorb photons of many different wavelengths, move to higher MO energy levels, and then return and release those photons explains metallic luster.
- In any period from group-1 to group-6 on the Periodic Table, the number of electrons in bonding MOs increase which will strengthen the bonds. However, from group-7 to group-12, the number of antibonding electrons begins to increase which weakens the bonds. Thus the observed trends for heats of fusion, hardness, and melting points can be explained.

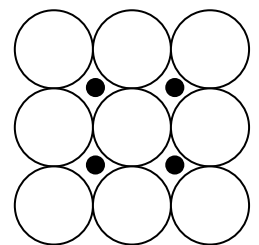
2) Metallic Alloys

- The properties of pure metals can be altered by mixing in other elements. This produces a material called an **alloy**.
- If the mixture formed is homogeneous, then the alloy is called a **solution alloy**. (Alloys can also be heterogeneous mixtures or intermetallic compounds.)
- There are two types of solution alloys:

☛ In **substitutional alloys**, solute atoms of similar radius replace some of the solvent atoms in the lattice structure. These alloys remain malleable and ductile and their densities are typically are between those of the constituent metals. **Example: brass – an alloy of copper with zinc or sterling silver – silver with copper**



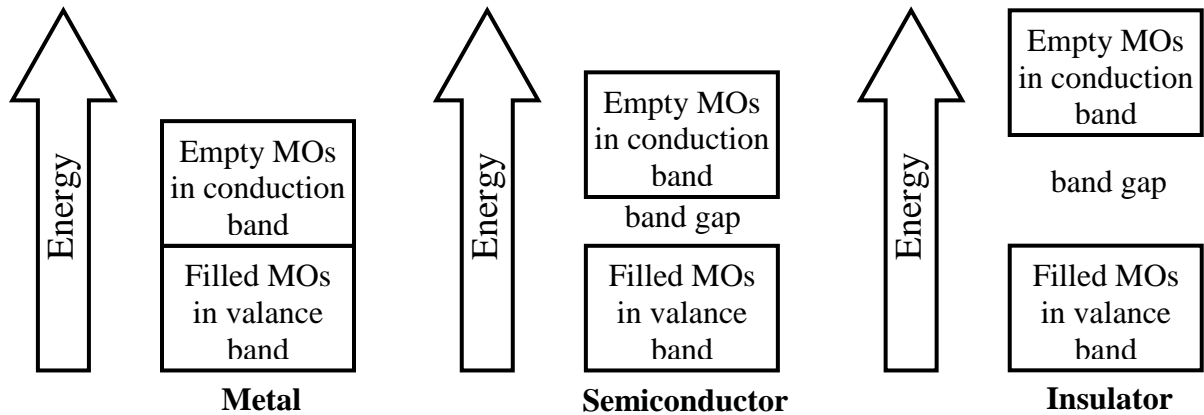
☛ In **interstitial alloys**, solute atoms of much smaller radius filled in empty spaces found between the solvent atoms in the lattice structure. These alloys have a more rigid lattice structure, making them less malleable and ductile. **Example: steel – an alloy of iron with carbon or stainless steel – an alloy of an alloy of iron with carbon and chromium.**



- In some cases, alloy formation changes the chemistry of the surface. Iron and steel reacts in air to form an oxide which does **not** adhere to the surface. Iron and steel rusts! Stainless steel reacts in air to form an inert oxide that strongly adheres to the surface forming a protective coating!

B) Semiconductors:

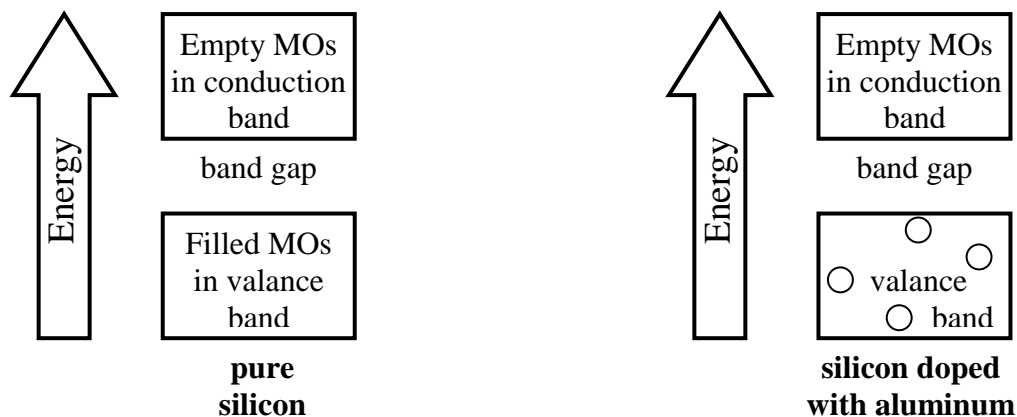
- A **semiconductor** is a material whose electrical conductivity is limited due to a gap exists between the filled and empty valence molecular orbitals. Only a small number of electrons can enter the conduction band at room temperature.



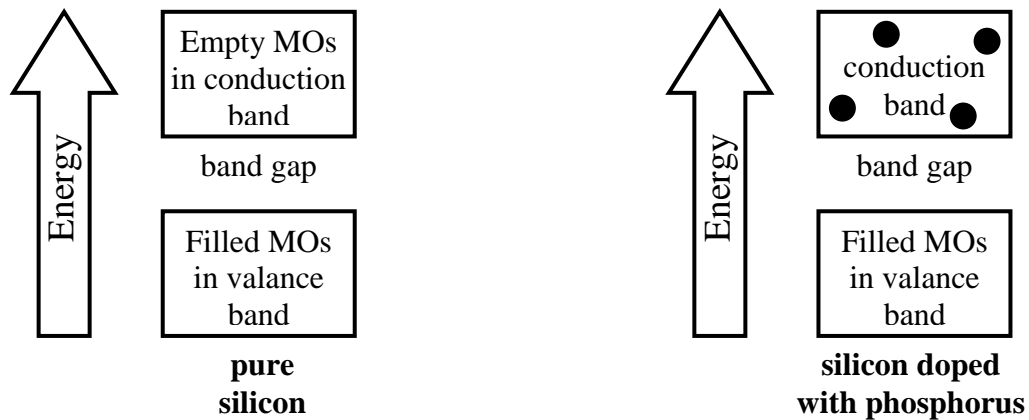
- Pure silicon has the same structure as diamond. Unlike diamond, an insulator, the **band (energy) gap** between the filled and empty valence molecular orbitals in silicon is small enough to let a few electrons cross at 25°C, making silicon a semiconductor.

1) The Doping of Semiconductors

- A semiconductor's conductivity can be increased by adding small amounts of an impurity on the order of just a few parts per million. This is called **doping**.
- When controlled amounts of aluminum are added to silicon, the result is called a **p-type semiconductor** because there are now fewer electrons present which form “vacancies” in the valence band of silicon. These “vacancies” are known as **holes** (the hole is thought of as having a positive charge).

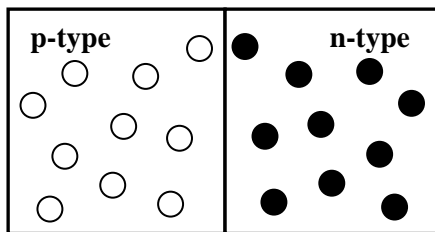


- When controlled amounts of phosphorus are added to silicon, the result is called an **n-type semiconductor** because there are now extra electrons present which must occupy the conduction band in silicon.

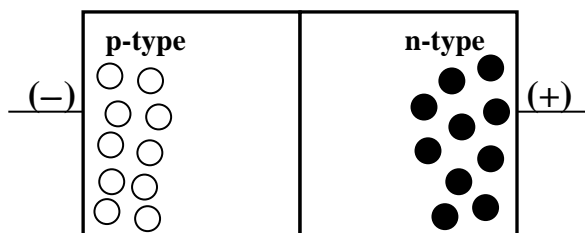


2) The Junction between p-Type and n-Type Semiconductors

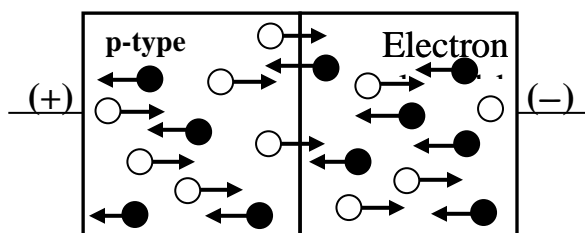
- When a p-type semiconductor is connected to an n-type semiconductor a **p-n junction** is formed and is used to control electron flow.



A p-n junction is formed when a p-type semiconductor is connected to an n-type semiconductor.



A p-n junction in which the p-type region is connected to the negative terminal and the n-type region is connected to the positive terminal. **In this case, there is NO electron flow across the p-n junction!**



A p-n junction in which the p-type region is connected to the positive terminal and the n-type region is connected to the negative terminal. **In this case, there IS electron flow across the p-n junction!**

- Such unidirectional current flow makes a p-n junction into a device that turns alternating current into direct current which is required for all electronic devices!

C) Glass and Other Ceramics:

- Glass and other ceramics are solid nonmetallic materials that have ionic and network covalent bonding present and have a network-like structure.

3) Glass

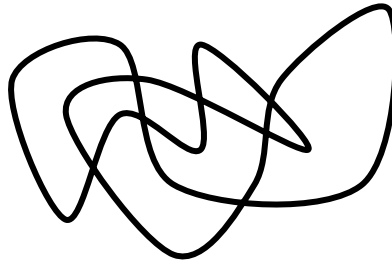
- Glass is a homogeneous noncrystalline material that is composed mostly of quartz – a silicate.
- When quartz melts ($\approx 1600^\circ\text{C}$) many of the Si-O bonds are broken a tacky liquid is produced. When the liquid is rapidly cooled, the Si-O bonds are reformed before they can rearrange themselves back into their regular crystal lattice structure. What is produced is an amorphous network solid called **silica or quartz glass**.
- Different types of glass are made by introducing impurities to the quartz as it melts.
 - ☛ Adding CaO and Na₂O produces soda-lime glass which is used in windows and bottles.
 - ☛ Adding K₂O instead of Na₂O produces a harder glass.
 - ☛ Adding PbO instead of CaO produces lead crystal glass.
 - ☛ Adding B₂O₃ produces glass that withstands thermal shock.
 - ☛ Adding CoO produces “cobalt” glass that has a blue color.
- Glass can usually be melted and reformed.

2) Other Ceramics

- Other ceramics are heterogeneous materials that are made mostly of clay – an aluminosilicate.
- When clay is fired at high temperatures, the microcrystals of clay become suspended in a glass-like cement.
- Aluminosilicate ceramics are strong, brittle and resistant to heat and chemical attack.
- Brick, porcelain, and clay pottery are example of ceramics.
- Ceramics are used as electrical insulators, dental crowns, thermal shields, floor tiles, as well as cookware & tableware.
- Unlike glass, other ceramics can not be melted and reformed.

D) Polymers:

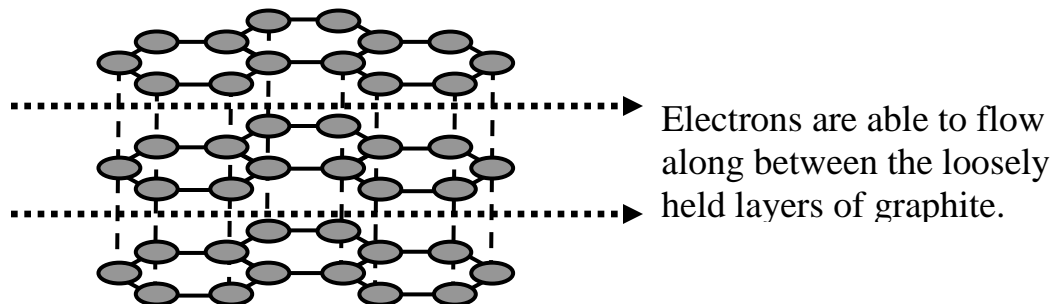
- A **polymer** is a material composed of many small molecules covalently bonded together into a very large molecule. This is called a **macromolecule** which is long chained structure.
- Proteins, plastics, rubbers, epoxy resins, nucleic acids, natural and synthetic fibers, cellulose, and starch are all polymers.
- The polymer's long chain can fold on itself and be held in place by either covalent bonds (cross-linking) or by intermolecular forces.



- The folding structure can be exact (crystalline) – as in proteins or random (amorphous) – as in many plastics.
- Once stretched, a plastic retains its new shape but a rubber will return to its original shape.

E) Graphite:

- The only pure elements that are semiconductors at 25°C are: (1) silicon, (2) germanium, and (3) carbon in its graphite form.
- Unlike silicon or germanium, graphite does not have a band gap.
- Graphite's semiconducting property is due to its structure.



- Graphite is also unusual in that it has a high melting point but is soft. It has a high melting point because of the strong covalent bonding that occurs between the carbon atoms in each layer. It is soft because the adjacent layers can easily slide past each other. The layers are attracted to each other by weak dispersion forces.